Chemistry 129.01 - General Chemistry Workshop - Spring 2017

Week #11

Monday, April 17. Buffer Solutions
Assigned reading: Section 14.6

Today we are going to begin our discussion of the important topic of buffers. Really, buffers are just equilibrium problems where you start the "reaction" with both reactants AND products and then solve for the equilibrium concentrations. In other words, since you know how to solve equilibrium problems there is really nothing new here. You simply need to solve the equilibrium problem with starting conditions that include both reactants and products.

We'll discuss talks the "Common-Ion Effect". This is just Le Chatelier's principle that we have talked about already. Make sure that you understand Le Chatelier's principle in this context, and remember the name "Common-Ion Effect" because it helps you recognize buffer type problems.

Buffers are cool because they are so important. Since all of the enzymes in our body require a specific pH to operate at maximum efficiency, the pH in our bodies must be held very constant. Blood is therefore *buffered* against changes in pH. Either strong acid OR strong base can be added to our blood without changing the pH very much. How is this done?

By having both the acid and the conjugate base of a weak acid in solution at the same time, a solution can be protected from large changes in pH. Think about what would happen if you added strong acid to a solution that has a weak base in it. The acid would react with the weak base rather than water. If you add a strong base to a solution with a weak acid in it, the strong base reacts with the weak acid rather than water. If neither the strong acid nor the strong base can react with water, there is little change in pH.

1. Give the chemical formula of:

   The conjugate acid of NH_3:

   The conjugate base of HNO_2:

2. Consider the following equilibrium:

$$HSO_4^-{}_{(aq)} + HCO_3^-{}_{(aq)} \Leftrightarrow SO_4^{2-}{}_{(aq)} + H_2CO_3{}_{(aq)}$$

   How will the system respond if more SO_4^{2-} were added to the solution? How about HCO_3^-?

Notes:

Chemistry 129.01 - General Chemistry Workshop - Spring 2017

Friday, April 21. Acid-Base Titrations
Assigned reading: Sections 14.7

Quiz today: Acid-Base Equilibria

In lab, we have looked at titrations as an analytical method. We slowly react two reagents' solutions (one of known concentration and the other of unknown concentration) until the equivalence point is reached (point at which stoichiometrically equivalent amounts have been mixed). Then, based on our results, we can determine the unknown concentration. Today we'll revisit acid-base titrations this time to track the changes in pH of the solution upon the addition of increasing volumes of titrant (solution delivered by the burette).

We'll look at titrations of strong acids and strong bases as well as titrations of weak acids and weak bases curves. We'll discuss the titration curves (plots of pH vs volume of titrant added) for each case highlighting the differences between them.

1. Define the following terms:

    a. Indicator

    b. Buffer solution

2. Write the balanced chemical formula for the reaction of the following:

    a. $HClO_{4\ (aq)}$ and $KOH_{\ (aq)}$

    b. $HCl_{\ (aq)}$ and $NH_{3\ (aq)}$

Notes:

Chemistry 129.01 - General Chemistry Workshop - Spring 2017

Problem Set #11

Due Monday, April 24 (at the beginning of class). Late homework will not be accepted.

1. Classify the following salts as acidic, basic or neutral.
   (a) $FeCl_3$
   (b) NaF
   (c) KCl
   (d) $(CH_3)_3NHCl$
   (e) $NH_4C_2H_3O_2$

2. In which of these solutions will HNO_2 ionize less than it does in pure water? Why?
   (a) 0.10M NaCl
   (b) 0.10M NaOH
   (c) 0.10M KNO_3
   (d) 0.10M $NaNO_2$

3. A buffer contains significant amounts of ammonia (NH_3) and ammonium chloride (NH_4Cl). Write equations showing how this buffer neutralizes added acid and added base.

4. A 1.00L buffer solution is 0.15M in hypobromous acid (HBrO) and 0.18M in sodium hypobromite (NaBrO).
   (a) What is the pH of this buffer?
   (b) What is the pH of the buffer after the addition of 0.01 mol of NaOH? (c) What is the pH of the buffer after the addition of 0.01 mol of HCl?

5. How many grams of sodium acetate ($NaC_2H_3O_2$) should be added to 1.00L of 0.150M acetic acid ($HC_2H_3O_2$) to form a buffer with pH 5.00? Assume that no volume change occurs when the sodium acetate is added.

6. Predict whether the pH of the equivalence point of each of the following titrations is below, above or at pH 7:
   (a) hydrofluoric acid (HF) titrated with KOH
   (b) perchloric acid ($HClO_4$) titrated with NaOH
   (c) methylamine (CH_3NH_2) titrated with HCl

7. A 25.0mL sample of 0.240M NaOH solution is titrated with a 0.200M HNO_3 solution. Calculate the pH of the solution after the following volumes of acid have been added: (a) 0.0mL, (b) 15.0mL, (c) 25.0mL, (d) 30.0mL, (e) 45.0mL.

8. A 25.0mL sample of 0.200M $HC_2H_3O_2$ solution is titrated with a 0.250M KOH solution. Calculate the pH of the solution after the following volumes of base have been added: (a) 0.0mL, (b) 10.0mL, (c) 15.0mL, (d) 20.0mL, (e) 30.0mL.

9. A sample of an unknown monoprotic acid was dissolved in distilled water and titrated with a NaOH solution. The pH of the solution was monitored throughout the titration, and the following data was collected. Determine the K_a of the acid. Hint: You can use Excel to plot the titration curve.

mL NaOH added	**pH**	**mL NaOH added**	**pH**
0	3.09	22.0	5.93
5	3.65	22.2	6.24
10	4.10	22.6	9.91
15	4.50	22.8	10.21
17	4.55	23.0	10.38
18	4.71	24.0	10.81
19	4.94	25.0	11.02
20	5.11	30.0	11.49
21	5.37	40.0	11.85